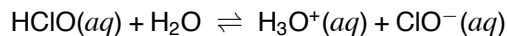
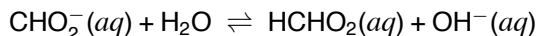
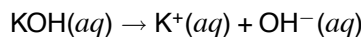
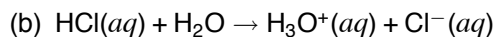


SUMMARY PROBLEM

(a) HCl – strong acid, KOH – strong base; HCHO₂ – weak base, HClO – weak acid



(c) HClO/ClO[−];

CHO₂[−]/HCHO₂

(d) HCl: pH = 0.67;

KOH: pH = 13.33

HClO: $2.8 \times 10^{-8} = \frac{(x)(x)}{0.215 - x};$

$x = [\text{H}^+] = 7.8 \times 10^{-5}$ pH = 4.11

KCHO₂: $5.3 \times 10^{-11} = \frac{(x)(x)}{0.215 - x};$

$x = [\text{OH}^-] = 3.4 \times 10^{-6};$ pH = 8.53

(e) $K_b \text{ for ClO}^- = \frac{1.0 \times 10^{-14}}{2.8 \times 10^{-8}} = 3.6 \times 10^{-7};$

$K_a \text{ for HCHO}_2 = \frac{1.0 \times 10^{-14}}{5.3 \times 10^{-11}} = 1.9 \times 10^{-4}$

(f) NH₄Cl – acidic, KClO₄ – neutral

(g) $K_a\text{NH}_4^+ = 5.6 \times 10^{-10};$ $K_b\text{ClO}^- = 3.6 \times 10^{-7};$ $K_b > K_a;$ salt is basic

PROBLEMS

1. (a) acids: H₃O⁺, HCN; bases: CN[−], H₂O; conjugate pairs: H₃O⁺, H₂O; HCN, CN[−]

(b) acids: HNO₂, H₂O; bases: NO₂[−], OH[−]; conjugate pairs: HNO₂, NO₂[−]; H₂O, OH[−]

(c) acids: HCHO₂, H₃O⁺; bases: CHO₂[−], H₂O; conjugate pairs: HCHO₂, CHO₂[−]; H₃O⁺, H₂O

3. (a) acid

(b) acid

(c) base

5. (a) H₂O

(b) H₂PO₄[−]

(c) NH₄⁺

(d) HF

(e) Zn(H₂O)₃(OH)⁺

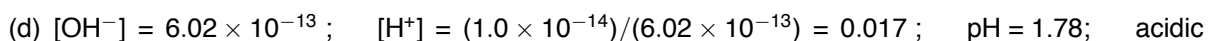
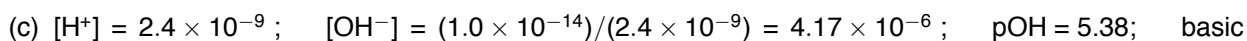
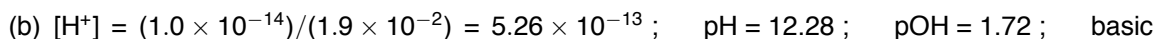
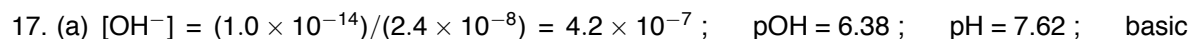
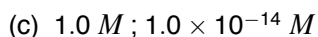
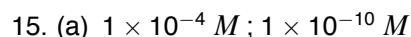
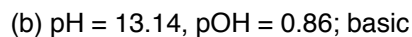
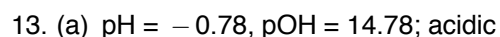
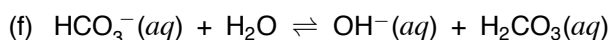
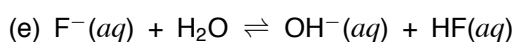
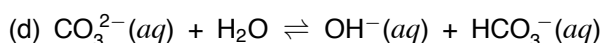
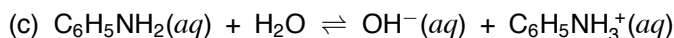
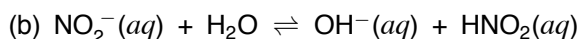
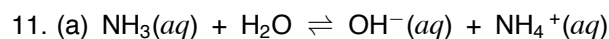
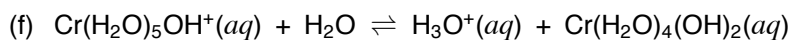
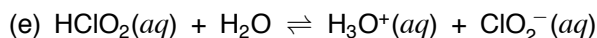
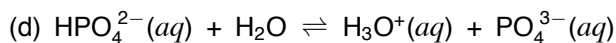
7. As B-L acid: $\text{H}_2\text{PO}_4^-(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HPO}_4^{2-}(aq)$

As B-L base: $\text{H}_2\text{PO}_4^-(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{PO}_4(aq) + \text{OH}^-(aq)$

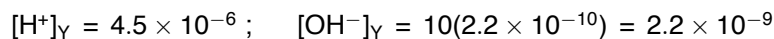
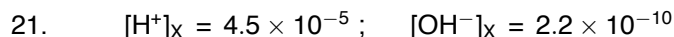
9. (a) $\text{Ni}(\text{H}_2\text{O})_5\text{OH}^+(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{Ni}(\text{H}_2\text{O})_4(\text{OH})_2(aq)$

(b) $\text{Al}(\text{H}_2\text{O})_6^{3+}(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{Al}(\text{H}_2\text{O})_5(\text{OH})^{2+}(aq)$

(c) $\text{H}_2\text{S}(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+(aq) + \text{HS}^-(aq)$



19. Solution 1 is more acidic. Solution 2 has the higher pH.



$$(a) [H^+]_X/[H^+]_Y = 4.5 \times 10^{-5}/4.5 \times 10^{-6} = 10; \quad [H^+]_X/[H^+]_Z = 4.5 \times 10^{-5}/4.5 \times 10^{-9} = 1.0 \times 10^4$$

$$(b) (pH)_Y = 5.35; (pH)_Z = 8.35$$

(c) X and Y are acidic, Z is basic.

$$23. \text{rain water: } [H^+] = 3 \times 10^{-6}; \quad \text{acid rain: } [H^+] = 1 \times 10^{-3}; \quad \text{ratio} = 1 \times 10^{-3}/3 \times 10^{-6} = 3 \times 10^2$$

$$25. (a) \text{mol OH}^- = 0.25 \text{ g Ba(OH)}_2 \times \frac{1 \text{ mol Ba(OH)}_2}{171.3 \text{ g Ba(OH)}_2} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ba(OH)}_2} = 2.9 \times 10^{-3}$$

$$[OH^-] = \frac{0.0029 \text{ mol}}{0.655 \text{ L}} = 0.0045 \text{ M}; \quad pOH = 2.35; \quad pH = 11.65$$

$$(b) \text{mol KOH} = (0.300 \text{ L})(0.149 \text{ mol/L}) = 0.0447; \quad M \text{ of diluted sol'n} = 0.0447 \text{ mol}/3.00 \text{ L} = 0.0149$$

$$pOH \text{ of concentrated solution} = 0.83; \quad pH \text{ of concentrated solution} = 13.17$$

$$pOH \text{ of diluted solution} = 1.83; \quad pH \text{ of diluted solution} = 12.17$$

pH decreases by 1.00 unit

$$27. (a) [OH^-] = 0.27 \text{ M Sr(OH)}_2 \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Sr(OH)}_2} = 0.54$$

$$[H^+] = 1.0 \times 10^{-14}/0.54 = 1.9 \times 10^{-14}; \quad pH = 13.73; \quad pOH = 0.27$$

$$(b) \text{mol OH}^- = \text{mol KOH} = (13.6 \text{ g})(1 \text{ mol}/56.1 \text{ g}) = 0.242; \quad [OH^-] = \frac{0.242 \text{ mol}}{2.50 \text{ L}} = 0.0970 \text{ M}$$

$$[H^+] = 1.00 \times 10^{-13}; \quad pH = 13.00; \quad pOH = 1.00$$

$$29. \text{mol H}^+ \text{ from HNO}_3 = \text{mol HNO}_3 = (0.786 \text{ mol/L})(0.295 \text{ L}) = 0.232$$

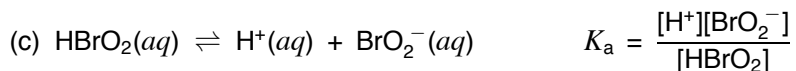
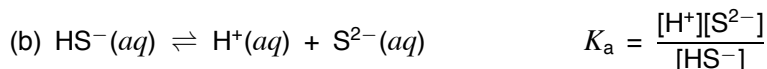
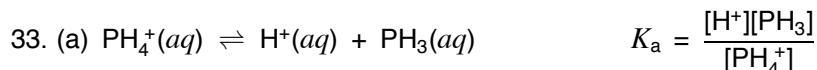
$$\text{mol H}^+ \text{ from HI} = \text{mol HI} = (5.00 \text{ g})(1 \text{ mol}/127.9 \text{ g}) = 0.0391$$

$$[H^+] = \frac{(0.232 + 0.0391) \text{ mol}}{0.295 \text{ L}} = 0.919 \text{ M}; \quad pH = 0.037$$

$$31. \text{mol OH}^- \text{ from NaOH} = \text{mol NaOH} = (13.0 \text{ g})(1 \text{ mol}/40.0 \text{ g}) = 0.325$$

$$\text{mol OH}^- \text{ from Sr(OH)}_2 = 2(\text{mol Sr(OH)}_2) = 2(0.200 \text{ mol/L})(0.795 \text{ L}) = 0.318$$

$$[OH^-] = \frac{(0.325 + 0.318) \text{ mol}}{0.795 \text{ L}} = 0.809 \text{ M}; \quad pOH = 0.0921; \quad pH = 13.91$$



$$35. (a) K_a \approx 1 \times 10^{-4} \quad (b) 7.6 \times 10^{-11} \quad (c) 8.5 \times 10^{-14}$$

$$37. (a) D > A > C > B \quad (b) D$$

$$39. \text{HClO} < \text{HC}_7\text{H}_5\text{O}_2 < \text{HNO}_2 < \text{HNO}_3$$

$$41. \text{HNO}_3 < \text{HNO}_2 < \text{HC}_7\text{H}_5\text{O}_2 < \text{HClO}$$

$$43. [\text{H}^+] = 4.6 \times 10^{-3} = [\text{B}^-]; \quad [\text{HB}] = 0.129 - 4.6 \times 10^{-3} = 0.124$$

$$K_a = \frac{(4.6 \times 10^{-3})(4.6 \times 10^{-3})}{0.124} = 1.7 \times 10^{-4}$$

$$45. M \text{ of caproic acid} = (0.450 \text{ mol})/2.0 \text{ L} = 0.225; \quad K_a = \frac{(1.7 \times 10^{-3})(1.7 \times 10^{-3})}{0.225 - (1.7 \times 10^{-3})} = 1.3 \times 10^{-5}$$

$$47. \text{mol HC}_6\text{H}_7\text{O}_6 (\text{HAsc}) = \frac{2.00 \text{ g}}{176.1 \text{ g/mol}} = 0.0114; \quad [\text{HAsc}]_0 = \frac{0.0114 \text{ mol}}{0.100 \text{ L}} = 0.114 \text{ M}; \quad [\text{H}^+]_{\text{eq}} = 0.0029 \text{ M}$$

$$K_a = \frac{(0.0029)(0.0029)}{0.114 - 0.0029} = 7.6 \times 10^{-5}$$

$$49. 1.2 \times 10^{-5} = \frac{[\text{H}^+][\text{Al}(\text{H}_2\text{O})_5\text{OH}^{2+}]}{1.75}; \quad [\text{H}^+] = [\text{Al}(\text{H}_2\text{O})_5\text{OH}^{2+}]; \quad [\text{H}^+] = 0.0046 \text{ M}; \quad \text{pH} = 2.34$$

$$51. (a) 1.51 \times 10^{-5} = \frac{[\text{H}^+]^2}{0.279 \text{ mol}/1.30 \text{ L}} \quad [\text{H}^+] = 1.80 \times 10^{-3} \text{ M}$$

$$(b) \text{mol butyric acid} = \frac{13.5 \text{ g}}{88.1 \text{ g/mol}} = 0.153; \quad M \text{ butyric acid} = \frac{0.153 \text{ mol}}{1.30 \text{ L}} = 0.118 \text{ mol/L}$$

$$1.51 \times 10^{-5} = \frac{[\text{H}^+]^2}{0.118} \quad [\text{H}^+] = 1.33 \times 10^{-3} \text{ M}$$

$$53. (a) \quad 5.1 \times 10^{-6} = \frac{[H^+]^2}{0.894 M} \quad [H^+] = 2.1 \times 10^{-3} M$$

$$(b) \quad [OH^-] = \frac{1.0 \times 10^{-14}}{2.1 \times 10^{-3}} = 4.8 \times 10^{-12} M$$

$$(c) \quad pH = 2.68$$

$$(d) \quad \% \text{ ionization} = \frac{2.1 \times 10^{-3}}{0.894} \times 100 = 0.23\%$$

$$55. (a) \quad [\text{phenol}] = \frac{(14.5/94.1) \text{ mol}}{0.892 \text{ L}} = 0.173 M; \quad [H^+]^2 = (1.1 \times 10^{-10})(0.173) = 1.9 \times 10^{-11}$$

$$[H^+] = 4.4 \times 10^{-6}; \quad pH = 5.36$$

$$(b) \quad \% \text{ ionization} = \frac{4.4 \times 10^{-6}}{0.173} \times 100 = 0.0025\%$$

$$57. \quad H_2C_2O_4(aq) \rightleftharpoons 2H^+(aq) + C_2O_4^{2-}(aq)$$

$$K = K_1 K_2 = (5.9 \times 10^{-2})(5.2 \times 10^{-5}) = 3.1 \times 10^{-6}$$

$$59. \quad 2.7 \times 10^{-4} = \frac{[H^+][HA^-]}{[H_2A]}; \quad [H^+] = [HA^-]; \quad [H^+]^2 = (2.7 \times 10^{-4})(0.20)$$

$$[H^+] = [HA^-] = 7.3 \times 10^{-3} M; \quad pH = 2.13; \quad [A^{2-}] = K_{a2} = 8.3 \times 10^{-7} M$$

$$61. \quad 1.2 \times 10^{-3} = \frac{[H^+][HC_8H_4O_4^-]}{[H_2C_8H_4O_4]}; \quad [H^+] = [HC_8H_4O_4^-]; \quad [H^+]^2 = (1.2 \times 10^{-3})(2.9)$$

$$[H^+] = [HC_8H_4O_4^-] = 5.9 \times 10^{-2} M; \quad pH = 1.23; \quad [C_8H_4O_4^{2-}] = K_{a2} = 3.9 \times 10^{-5} M$$

$$63. (a) \quad F^-(aq) + H_2O \rightleftharpoons HF(aq) + OH^-(aq) \quad K_b = \frac{[HF][OH^-]}{[F^-]}$$

$$(b) \quad HCO_3^-(aq) + H_2O \rightleftharpoons H_2CO_3(aq) + OH^-(aq) \quad K_b = \frac{[H_2CO_3][OH^-]}{[HCO_3^-]}$$

$$(c) \quad CN^-(aq) + H_2O \rightleftharpoons HCN(aq) + OH^-(aq) \quad K_b = \frac{[HCN][OH^-]}{[CN^-]}$$

$$65. \quad H_2PO_4^- < NO_2^- < CO_3^{2-}$$

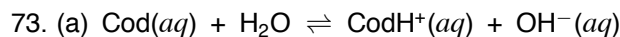
$$67. \quad (b) < (c) < (a) < (d)$$

69. (a) $K_b = 1.0 \times 10^{-14} / 0.16 = 6.2 \times 10^{-14}$

(b) $K_b = 1.0 \times 10^{-14} / 0.20 = 5.0 \times 10^{-14}$

71. $K_b = \frac{[\text{OH}^-][\text{HCO}_3^-]}{[\text{CO}_3^{2-}]}$; $[\text{OH}^-] = [\text{HCO}_3^-]$; $[\text{OH}^-]^2 = (2.1 \times 10^{-4})(0.28)$

$[\text{OH}^-] = 0.0077$; $\text{pOH} = 2.11$; $\text{pH} = 11.88$



(b) $K_b = \frac{1.0 \times 10^{-14}}{1.2 \times 10^{-8}} = 8.3 \times 10^{-7}$

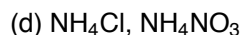
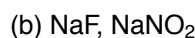
(c) $K_b = \frac{[\text{OH}^-][\text{CodH}^+]}{[\text{Cod}]}$; $[\text{OH}^-] = [\text{CodH}^+]$; $1.7 \times 10^{-9} = \frac{[\text{OH}^-]^2}{2.0 \times 10^{-3}}$

$[\text{OH}^-] = 4.1 \times 10^{-5} M$; $[\text{H}^+] = 2.4 \times 10^{-10} M$; $\text{pH} = 9.62$

75. $\text{pOH} = 14.00 - 11.68 = 2.32$; $[\text{OH}^-] = 4.8 \times 10^{-3} M$

$1.8 \times 10^{-5} = \frac{(4.8 \times 10^{-3})(4.8 \times 10^{-3})}{[\text{NH}_3]}$; $[\text{NH}_3] = 1.3 M$

$\text{mol NH}_3 = (1.3 \text{ mol/L})(1.25 \text{ L}) = 1.6$; $\text{mass NH}_3 = (1.6 \text{ mol})(17.0 \text{ g/mol}) = 27 \text{ g}$



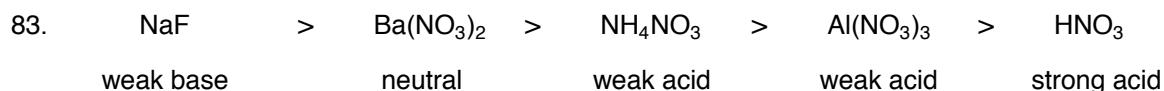
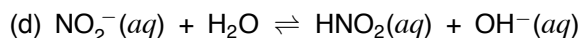
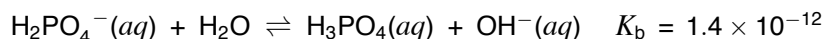
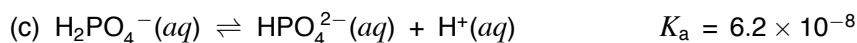
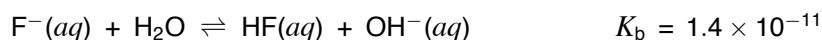
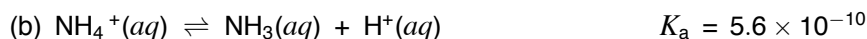
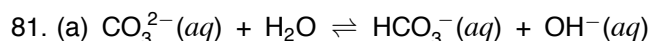
79. (a) basic

(b) slightly acidic

(c) acidic

(d) basic

(e) neutral



$$85. [\text{OH}^-] = (0.20 \text{ M})(0.05) = 0.010 \text{ M}; \quad K_b = \frac{(0.010)^2}{0.20 - 0.010} = 5.3 \times 10^{-4}$$

$$87. \text{mol aspirin} = \frac{0.648 \text{ g}}{180.15 \text{ g/mol}} = 3.60 \times 10^{-3}; \quad V = \frac{1}{16} \text{ qt} \times \frac{1 \text{ L}}{1.057 \text{ qt}} = 0.0591 \text{ L}$$

$$[\text{aspirin}] = 3.60 \times 10^{-3} \text{ mol} / 0.0591 \text{ L} = 0.0609 \text{ M}$$

$$[\text{H}^+]^2 = (3.6 \times 10^{-4})(0.0609 - [\text{H}^+]); \quad \text{The assumption that } \text{H}^+ \ll 0.0609 \text{ is not valid.}$$

Using the quadratic equation or the method of successive approximations, $[\text{H}^+] = 0.0045 \text{ M}$; $\text{pH} = 2.35$

$$89. \text{Assume } 100.0 \text{ g of solution.} \quad V_{\text{sol'n}} = \frac{100.0 \text{ g}}{1.10 \text{ g/mL}} = 90.9 \text{ mL}; \quad \text{mass of HY} = (100)(0.330) = 33.0 \text{ g}$$

$$\text{mol HY} = \frac{33.0 \text{ g}}{100 \text{ g/mol}} = 0.330; \quad [\text{HY}] = 0.330 \text{ mol} / 0.0909 \text{ L} = 3.63 \text{ M}$$

$$2.0 \times 10^{-8} = \frac{[\text{H}^+]^2}{3.63}; \quad [\text{H}^+] = 2.7 \times 10^{-4} \text{ M}; \quad \text{pH} = 3.57$$

91. No, K_W will be larger - endothermic reaction

$$\ln \frac{K_W \text{ at } 37^\circ\text{C}}{1.0 \times 10^{-14}} = \frac{55.8 \text{ kJ/mol}}{0.00831 \text{ kJ/mol} \cdot \text{K}} \left(\frac{1}{298 \text{ K}} - \frac{1}{310 \text{ K}} \right); \quad K_W \text{ at } 37^\circ\text{C} = 2.4 \times 10^{-14}$$

$$[\text{H}^+]^2 = 2.4 \times 10^{-14}; \quad [\text{H}^+] = 1.5 \times 10^{-7}; \quad \text{pH} = 6.81$$

$$93. \text{mol OH}^- = 2(\text{mol Mg(OH)}_2) = 2 \left(\frac{1.2 \text{ g}}{58.31 \text{ g/mol}} \right) = 0.0412$$

$$[\text{H}^+] = 0.032 \text{ M}; \quad \text{mol H}^+ = (0.032 \text{ mol/L})(0.200 \text{ L}) = 0.0063 \text{ mol need to be neutralized}$$

Since there are 0.0412 mol of OH^- , there is enough Mg(OH)_2 to neutralize the stomach acid.

95. neutral; $[\text{H}^+] = [\text{OH}^-]$

97. (a) T (b) F (c) T (d) T (e) T

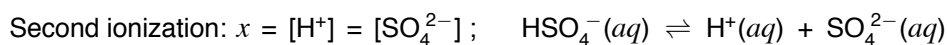
99. 1: 2 H^+ , 3 HX $K_a = 4/3 = 1.3$ strongest acid

2: 3 H^+ , 10 HX $K_a = 9/10 = 0.9$ smallest K_a , lowest pH

3: 2 H^+ , 4 HX $K_a = 4/4 = 1.0$

101. 1 H^+ , 1 A^- , 9 HA

103. H^+ from first ionization = $0.020M$



$$0.010 = \frac{(0.020 + x)(x)}{0.020 - x}; \quad x = 0.0056; \quad \text{H}^+_{\text{tot}} = 0.020 + 0.0056 = 0.0256 \quad \text{pH} = 1.59$$

104. (a) $\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O} \quad \Delta H^\circ = -55.8 \text{ kJ}$

$$\Delta H = 1.00 \text{ L} \times 0.100 \frac{\text{mol}}{\text{L}} \times \frac{-55.8 \text{ kJ}}{1 \text{ mol}} = -5.58 \text{ kJ}$$

(b) $\text{HF}(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O} + \text{F}^-(aq) \quad \Delta H^\circ = -68.3 \text{ kJ}$

$$\Delta H = 1.00 \text{ L} \times 0.100 \frac{\text{mol}}{\text{L}} \times \frac{-68.3 \text{ kJ}}{1 \text{ mol}} = -6.83 \text{ kJ}$$

105. $\text{mol H}^+ = (0.45 \text{ L})(0.12 \text{ mol/L}) = 0.054$

$$\text{mol OH}^- = 2(\text{mol Ba(OH)}_2) = 2(0.30 \text{ L})(0.233 \text{ mol/L}) = 0.14$$

$$\text{mol OH}^- \text{ not neutralized} = 0.14 - 0.054 = 0.086$$

$$[\text{OH}^-] = \frac{8.6 \times 10^{-2} \text{ mol}}{(0.30 + 0.45) \text{ L}} = 0.11 \text{ M}; \quad \text{pOH} = 0.94; \quad \text{pH} = 13.06$$

106. % ionization = $\frac{[\text{H}^+]}{[\text{HA}]_0} \times 100$; $[\text{H}^+]^2 = K_a \times [\text{HA}]_0$; $[\text{H}^+] = (K_a)^{\frac{1}{2}} \times ([\text{HA}]_0)^{\frac{1}{2}}$

$$\% \text{ ionization} = \frac{(K_a)^{\frac{1}{2}}}{([\text{HA}]_0)^{\frac{1}{2}}} \times 100$$

107. Consider 1000.0 g of solution: 950.0 g H_2O and 50.0 g $\text{HC}_2\text{H}_3\text{O}_2$

$$\text{mol HC}_2\text{H}_3\text{O}_2 = \frac{50.0 \text{ g}}{60.05 \text{ g/mol}} = 0.833$$

$$\text{before ionization: } m = \frac{0.838 \text{ mol}}{0.950 \text{ kg}} = 0.877$$

$$\text{volume of solution} = \frac{1000.0 \text{ g}}{1.006 \text{ g/mL}} = 994 \text{ mL}$$

$$[\text{HC}_2\text{H}_3\text{O}_2] = 0.833 \text{ mol}/0.994 \text{ L} = 0.838 \text{ M}; \quad [\text{H}^+] = \left[(0.838)(1.8 \times 10^{-5}) \right]^{\frac{1}{2}} = 0.0039 \text{ M}$$

$$\text{For } \text{H}^+ \text{ } M \approx m; \quad \text{total molality} = 0.0039 + 0.0039 + (0.877 - 0.0039) = 0.881 \text{ m}$$

$$\Delta T_f = (1.86^\circ\text{C}/m)(0.881 \text{ m}) = 1.64^\circ\text{C}; \quad T_f = -1.64^\circ\text{C}$$

108. molarity of $\text{Ca}(\text{OH})_2$: $M = \frac{(1.53/74.10) \text{ mol}}{1.00 \text{ L}} = 0.0206$

$$[\text{OH}^-] = 2[\text{Ca}(\text{OH})_2] = 2(0.0206) = 0.0412 \text{ M}; \quad \text{pOH} = 1.39; \quad \text{pH} = 12.61$$

109. $\text{mol HA} = \frac{11 \text{ g}}{138 \text{ g/mol}} = 0.080$; $[\text{HA}] = 0.080 \text{ mol}/0.745 \text{ L} = 0.107 \text{ M}$

$$\text{mol HB} = \frac{5.00 \text{ g}}{72.0 \text{ g/mol}} = 0.0694; \quad [\text{HB}] = 0.0694 \text{ mol}/0.525 \text{ L} = 0.132 \text{ M}$$

$$K_a \text{ HA} = [\text{H}^+]^2/0.107; \quad [\text{H}^+]^2 = (0.107)K_a \text{ HA}$$

$$K_a \text{ HB} = [\text{H}^+]^2/0.132; \quad [\text{H}^+]^2 = (0.132)K_a \text{ HB}$$

$$(0.107)K_a \text{ HA} = (0.132)K_a \text{ HB}$$

$$\text{Assume } K_a \text{ HB} = 1; \quad K_a \text{ HA} = \frac{(0.132)(1)}{0.107} = 1.23$$

$$K_a \text{ HA} > K_a \text{ HB}; \quad \text{HA is stronger.}$$